

Redox Reactions

OXIDATION AND REDUCTION

11.1

At the end of this worksheet you should know

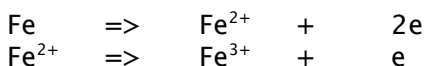
1. An oxidising agent is a substance, which accepts electrons; a reducing agent is a substance, which donates electrons.
2. Oxidising and reducing agents can be identified in redox reactions.

OXIDATION

Oxidation is defined as a loss of electrons.

e.g. iron rusting

When iron metal is exposed to oxygen in the air the iron combines with the oxygen to form iron (II) oxide and further reacts to form iron(III) oxide



REDUCTION

Reduction is defined as a gain of electrons.

e.g. Formation of copper from copper (II) oxide

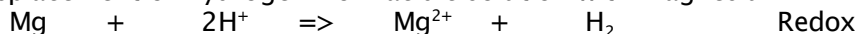
When hydrogen gas is passed over heated copper (II) oxide, the copper (II) oxide is reduced to copper metal.



Redox

Redox reactions are a combination of reduction and oxidation, which occur simultaneously.

e.g. Displacement of hydrogen from acidic solution with magnesium.



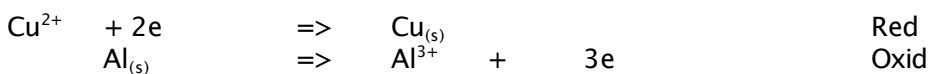
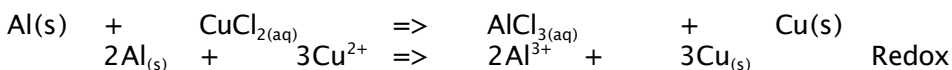
Red

Oxid

OXIDISING AND REDUCING AGENTS

An oxidising agent is an electron acceptor. An oxidising agent is **reduced** during a redox reaction.

A reducing agent is an electron donor. A reducing agent is **oxidised** during a redox reaction.



The reducing agent is The oxidising agent is

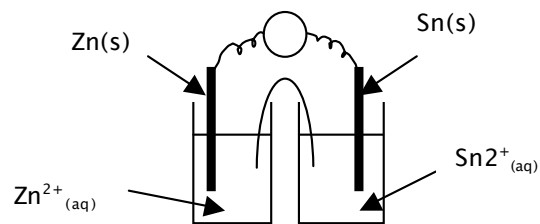
Note The spectator ions in the above reaction were

Consider the cell opposite



Red

Oxid



The reducing agent is The oxidising agent is

Redox equations

11.2

At the end of this worksheet you should know that

1. Ion-electron equations can be written for oxidation and reduction reactions.

- Ion-electron equations can be combined to give redox equations
- Given reactant and product species, ion electron equations, which include H^+ (aq), and $\text{H}_2\text{O}_{(l)}$ can be written.

WRITING A BALANCED REDOX EQUATION

In order to write a balanced redox equation for a reaction we have to

- write the ion-electron equation for the oxidation.
- write the ion-electron equation for the reduction.
- balance the number of electrons gained and lost.
- add the oxidation and reduction reactions together.

e.g. when magnesium is added to a solution of iron(III) nitrate the magnesium displaces the iron(III) ions from solution to form iron and magnesium nitrate.

The ion-electron equation for the oxidation is:-

The ion-electron equation for the reduction is:-

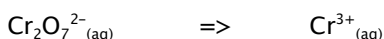
To balance the number of electrons gained and lost we must multiply the oxidation reaction by 3 and the reduction reaction by 2 then add them together giving:-

In the examples we have looked at so far the oxidation and reduction reactions have always been relatively easy to write in balanced form. However consider the reaction between acidified potassium dichromate solution and tin(II) chloride solution in which tin(II) ions are oxidised to tin(IV) ions and dichromate ions reduced to chromium(III) ions.

The ion-electron equation for the oxidation is straightforward:-

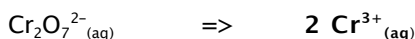


The reduction of the dichromate ions to chromium(III) ions as shown below is **not** a balanced equation.



In order to balance this equation we have to apply the following rules:-

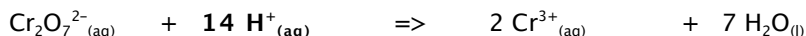
- Balance the number of chromium atoms on both sides of the reaction.



- Add sufficient water molecules to balance the number of oxygen atoms.



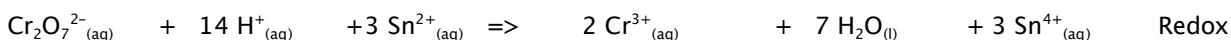
- Add sufficient hydrogen ions to balance the hydrogen in the water molecules.



- Add the correct number of electrons to balance the total charge on each side.

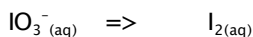


The above ion-electron equation is now balanced and shows the reduction of dichromate to chromium(III) ions.

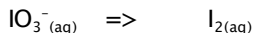


In acid solution, iodate ions, $\text{IO}_3^-_{(aq)}$, are readily converted into iodine (I_2). Write an ion-electron equation for the reaction.

- Balance the number of iodine atoms on both sides of the reaction.



- Add sufficient water molecules to balance the number of oxygen atoms.



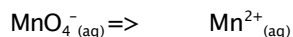
- Add sufficient hydrogen ions to balance the hydrogen in the water molecules.

4. Add the correct number of electrons to balance the total charge on each side.

The above ion–electron equation is now balanced and shows the reduction of iodate to iodine molecules.

In acid solution permanganate ions MnO_4^- are converted to manganese(II) ions Mn^{2+}

1. The number of manganese atoms on both sides of the reaction are already balanced.



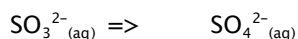
2. Add sufficient water molecules to balance the number of oxygen atoms.

3. Add sufficient hydrogen ions to balance the hydrogen in the water molecules.

4. Add the correct number of electrons to balance the total charge on each side.

Sulphite ions SO_3^{2-} act as a reducing agent, changing to sulphate ions SO_4^{2-} .

1. The number of sulphur atoms on both sides of the reaction is already balanced.

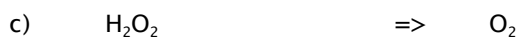
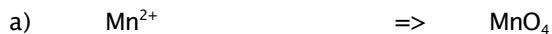


2. Add sufficient water molecules to balance the number of oxygen atoms.

3. Add sufficient hydrogen ions to balance the hydrogen in the water molecules.

4. Add the correct number of electrons to balance the total charge on each side.

Complete each of the following



REDOX TITRATIONS

11.3

At the end of this worksheet you should know

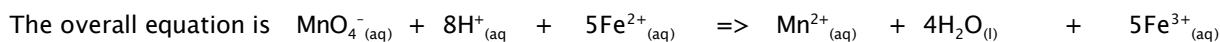
1. The concentration of a reactant can be determined from the results of a redox titration.

A redox titration is not unlike an acid-alkali titration but the reactants are an oxidising agent and a reducing agent. The oxidising agent in this reaction is acidified potassium permanganate solution and the reducing agent is a solution of iron(II) sulphate.

The reduction reaction is :-



The oxidation reaction is:-

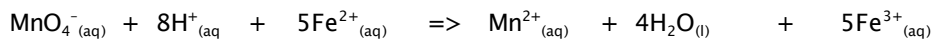


Note The purple permanganate ion reacts with the iron(II) ions to form the colourless manganese(II) ion. When all of the iron(II) ions in the solution have been reacted the next drop of permanganate added has nothing to react with and so the purple colour remains. Hence no indicator needed in this experiment as the permanganate acts as its own indicator.

Method A burette was rinsed and filled with standard 0.02M potassium permanganate solution. 20 cm³ of unknown iron(II) sulphate was pipetted into a clean conical flask. 10 cm³ of 1M sulphuric acid was added to the conical flask and titrated with the standard permanganate solution until a permanent purple colour was produced. The titration was repeated until concordant results were obtained.

Results	Volume of permanganate at start (cm ³)	=	2.0	16.2	30.8
	Volume of permanganate at end (cm ³)	=	16.2	30.8	45.2
	Volume of permanganate solution used (cm ³)	=	14.2	14.6	14.4

$$\text{Average titre} = (14.6 + 14.4)/2 = 14.5 \text{ cm}^3$$



$$\begin{array}{ll} V_x = 14.5 \text{ cm}^3 & V_y = 20 \text{ cm}^3 \\ C_x = 0.02\text{M} & C_y = ? \\ n_x = 1 & n_y = 5 \end{array}$$

To calculate the concentration of the iron(II) sulphate solution used we use the relationship:-

$$\frac{V_x \times C_x}{n_x} = \frac{V_y \times C_y}{n_y}$$

$$\frac{14.5 \times 0.02}{1} = \frac{20 \times C_y}{5}$$

$$\begin{aligned} C_y &= \frac{14.5 \times 0.02 \times 5}{20} \\ &= 0.0725 \text{ mol l}^{-1} \end{aligned}$$

Where

V_x is the volume of permanganate

C_x is the concentration of permanganate

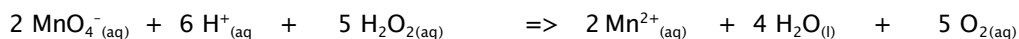
n_x is the number of moles of permanganate

V_y is the volume of iron(II) ions

C_y is the concentration of iron(II) ions

n_y is the number of moles of iron(II) ions

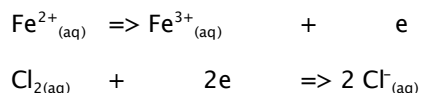
1. The concentration of hydrogen peroxide solution can be found by a redox titration with acidified potassium permanganate solution. It was found that 10 cm³ of the hydrogen peroxide solution reacted with 20 cm³ of 0.02M potassium permanganate solution when titrated.. Calculate the concentration of the hydrogen peroxide solution.



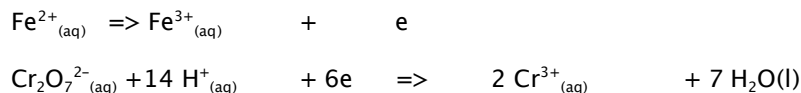
$$\begin{array}{ll} V_x = 20 \text{ cm}^3 & V_y = 10 \text{ cm}^3 \\ C_x = 0.02\text{M} & C_y = ? \\ n_x = 2 & n_y = 5 \end{array}$$

$$\frac{V_x \times C_x}{n_x} = \frac{V_y \times C_y}{n_y}$$

2. The chlorine levels in swimming pools can be determined by titrating samples against acidified iron (II) sulphate solution.. A 100 cm³ sample of water required 25 cm³ of 2.0M iron(II) sulphate to reach an end point. Calculate the concentration of the chlorine in the swimming pool. The ion –electron equations for the reaction are



3. Iron (II) ions react with dichromate ions in acidic solution. The ion electron equations being:-



If 20 cm³ of dichromate solution react with 60 cm³ of 0.2M iron(II) ions concentration what is the concentration of the dichromate solution?

4. Iodine solution (I_{2(aq)}) is decolourised by a solution containing thiosulphate ions (S₂O₃²⁻_(aq)), the iodine changing to iodide ions (I⁻_(aq)) and the thiosulphate changing to S₂O₆²⁻_(aq).

- Write ion – electron equations for oxidation and reduction reactions. (Note I₂)
- Write the redox reaction

Given that 20 cm³ of iodine solution were decolourised by 24.2 cm³ of 0.2 mol l⁻¹ thiosulphate solution calculate the concentration of the iodine solution in

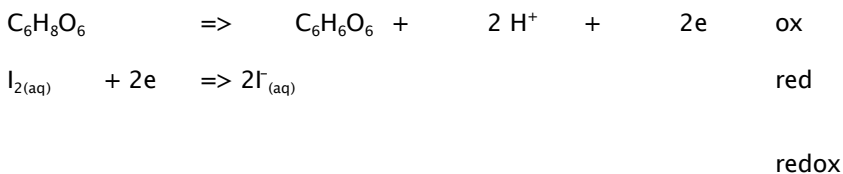
- mol l⁻¹ . (M)
- g l⁻¹.

Estimating Vitamin C

11.4

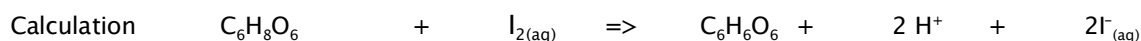
At the end of this unit you should be able to determine the mass of vitamin C in a vitamin C tablet.

Vitamin C is also called ascorbic acid, it is found naturally in many fruits and vegetables. It is an important component of the human diet. In its absence, a nasty skin condition called scurvy can result. Ascorbic acid can be dissolved in water and reacted with iodine solution.



Method A vitamin C tablets was crushed in a clean dry beaker. 50 cm³ of distilled water was added to the beaker and stirred to dissolve the tablet. The dissolved tablet was added to a graduated flask, the beaker was washed with two further aliquots of distilled water, which were also added to the graduated flask. The flask was filled to 250 cm³ with distilled water and carefully shaken to ensure proper mixing of the vitamin C solution. 25 cm³ of this solution was pipetted into a clean conical flask. This was titrated using standard 0.025 M iodine solution using starch solution as an indicator.

Results	Rough	1st	2nd
Burette reading at start of titration (cm ³)			
Burette reading at end of titration (cm ³)			
Volume of iodine solution used (cm ³)			
Average volume of titre (cm ³)			



$$n_x = 1$$

$$n_y = 1$$

n_x = moles of vit C in eq.

$$V_x = 25 \text{ cm}^3$$

$$V_y =$$

V_x = volume of vit C
 M_x = molarity of vit C

$$M_x = ?$$

$$M_y = 0.025$$

n_y = moles of iodine in eq.
 V_y = volume of iodine
 M_y = molarity of iodine

$$\frac{M_x V_x}{M_y V_y} = \frac{n_x}{n_y}$$

Use values to determine M_x

Use expressions $n = MV/1000$ and $n = \text{mass}/\text{GFM}$ to determine the mass of vitamin C in 1 tablet.

Electrolytic mass

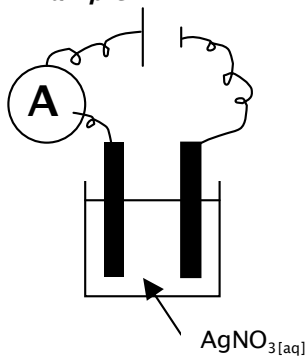
11.5

At the end of this unit you should know that

1. The production of one mole of an element from its ion, by electrolysis, always requires 'n' times 96,500 coulombs (C) where n is the number of electrons from the relevant ion-electron equation.
2. 96,500 C is the charge associated with one mole of electrons.
3. The mass of an element can be calculated from the quantity of electricity passes and vice-versa.
4. $Q = It$ where Q is quantity of charge in coulombs, I is the current in amps and t is time in seconds.

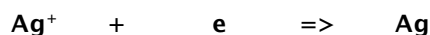
Electrolysis, the breaking down of compounds into elements by passing an electric current, is an example of a redox reaction

Example



What mass of silver would be deposited if 2A were passed for 30 mins?

$$Q = It = 2 \times 30 \times 60 = 3600 \text{ C}$$



$$96,500 \text{ C} \Rightarrow 108 \text{ g}$$

$$1 \text{ C} = 108/96,500$$

$$3600 \text{ C} = [108/96,500] \times 3600 = 4 \text{ g}$$

1. What mass of copper metal would be deposited by electrolysis copper(II) sulphate solution using 5A for 30 mins?
2. What mass of nickel would be deposited by passing 10A for 5 mins through a solution of nickel(II) nitrate?
3. A silver(I) nitrate solution was electrolysed. A steady current was maintained for 1 hour. If 2.16 g of silver was deposited, what current was used?
4. Calculate the mass of aluminium in kg produced by passing 175,000 A for 1 hour.
5. From the information below calculate the current used when nickel plating a sheet of metal.

$$\text{Mass of metal before plating} = 88.96 \text{ g}$$

$$\text{Mass of metal after plating} = 90.50 \text{ g}$$

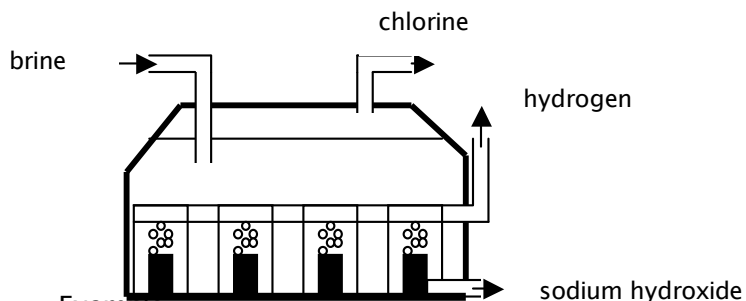
$$\text{Time taken} = 20 \text{ mins}$$

Gases by Electrolysis 11.6

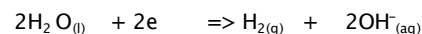
At the end of this unit you should be able to

1. State that the molar volume of any gas is constant at any given temperature and pressure.
2. Calculate the volume of a gas given the number of moles and vice – versa.
3. Calculate the volume of a gas produced by passing a known quantity of electricity.

Fluorine and chlorine are gases , which are usually produced by electrolysis as they are difficult to produce in other ways. Hydrogen and oxygen gases are often produced in electrolysis as by – products of other reactions.



At negative electrode



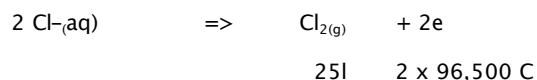
At positive electrode



Example

- a) If a current of 60,000 A is passed for 60 s what volume of chlorine will be evolved.
- b) If 40,000 A is passed for 5 mins , what volume of hydrogen gas will be produced (Molar volume = 25 l mol⁻¹.)

$$\text{a) } Q = It = 60,000 \times 60 = 3,600,000 \text{ C or } 3.6 \times 10^6$$



$$2 \times 96,500 \text{ C} = 25 \text{ l}$$

$$1 \text{ C} = 25 / 193,000$$

$$3,600,000 \text{ C} = 25 / 193,000 \times 3,600,000 \text{ C}$$

$$= 466.3 \text{ l}$$

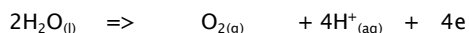
$$\text{b) } Q = It = 40,000 \times 5 \times 60 =$$



$$2 \times 96,500 \text{ C} = 25 \text{ l}$$

$$1 \text{ C} = 25 / 193,000$$

1. What volume of oxygen gas will be evolved when 10A is passed through dilute sulphuric acid for 10 mins. Molar volume is 24 l mol⁻¹.



2. What volume of hydrogen will be evolved when 20A is passed through dilute hydrochloric acid for 10 mins. Molar volume is 24 l mol⁻¹.

Quantitative Electrolysis 11.7

At the end of this unit you should be able to

1. Carry out an experiment to confirm that 2 moles of electrons i.e. 193,000 C are needed to

